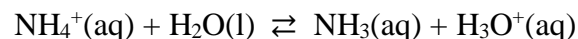


More examples of Buffer calculations

0.15 mol L⁻¹ NH₃ and 0.50 mol L⁻¹ NH₄Cl make up a buffer solution. What is the pH of the buffer solution? $K_a \text{ NH}_4^+ = 5.8 \times 10^{-10}$



Using K_a

$$K_a = \frac{[\text{NH}_3][\text{H}_3\text{O}^+]}{[\text{NH}_4^+]}$$

$$5.8 \times 10^{-10} = \frac{[0.15][\text{H}_3\text{O}^+]}{[0.50]}$$

$$\frac{(5.8 \times 10^{-10})(0.50)}{0.15} = [\text{H}_3\text{O}^+]$$

$$1.933 \times 10^{-9} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log 1.933 \times 10^{-9}$$

$$\text{pH} = 8.713$$

Answer: 8.71

OR

alternative method using H-H equation

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log (5.8 \times 10^{-10})$$

$$\text{p}K_a = 9.236$$

$$\text{pH} = 9.236 + \log \frac{0.15}{0.50}$$

$$\text{pH} = 9.236 + \log 0.3$$

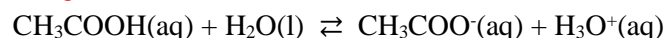
$$\text{pH} = 9.236 + (-0.5228)$$

$$\text{pH} = 8.713$$

Answer: 8.71

5.78g of sodium ethanoate was added to 1 litre of 0.1 mol L⁻¹ ethanoic acid. What is the pH of the buffer solution? $K_a (\text{CH}_3\text{COOH}) = 1.8 \times 10^{-5}$
 $M(\text{CH}_3\text{COONa}) = 82 \text{ g mol}^{-1}$

Assume that the added mass of sodium ethanoate does not change the total volume of solution



$$n = \frac{m}{M}$$

$$n(\text{CH}_3\text{COONa}) = \frac{5.78}{82}$$

$$n(\text{CH}_3\text{COONa}) = 0.07048 \text{ mol}$$

$$C(\text{CH}_3\text{COONa}) = \frac{n}{V}$$

$$C(\text{CH}_3\text{COONa}) = \frac{0.07048}{0.1} = 0.7048 \text{ mol L}^{-1}$$

Using K_a

$$K_a = \frac{[\text{CH}_3\text{COO}^-][\text{H}_3\text{O}^+]}{[\text{CH}_3\text{COOH}]}$$

$$1.8 \times 10^{-5} = \frac{0.7048 [\text{H}_3\text{O}^+]}{0.1}$$

$$2.553 \times 10^{-6} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log 2.553 \times 10^{-6}$$

$$\text{pH} = 5.592$$

Answer: 5.59

OR

alternative method using H-H equation

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log (1.8 \times 10^{-5})$$

$$\text{p}K_a = 4.744$$

$$\text{pH} = 4.744 + \log \frac{0.7048}{0.1}$$

$$\text{pH} = 4.744 + \log 7.048$$

$$\text{pH} = 4.744 + 0.8480$$

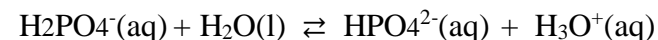
$$\text{pH} = 5.592$$

Answer: 5.59

2.64g of Na₂HPO₄·12H₂O and 0.31g of NaH₂PO₄·2H₂O made up a buffer solution with a volume of 1 litre. What is the pH of the buffer solution? $K_a \text{ H}_2\text{PO}_4^- = 6.3 \times 10^{-8}$

$$M(\text{Na}_2\text{HPO}_4 \cdot 12\text{H}_2\text{O}) = 358 \text{ g mol}^{-1}$$

$$M(\text{NaH}_2\text{PO}_4 \cdot 2\text{H}_2\text{O}) = 156 \text{ g mol}^{-1}$$



$$n = \frac{m}{M}$$

$n(\text{Na}_2\text{HPO}_4 \cdot 12\text{H}_2\text{O}) = \frac{2.64}{358}$	$n(\text{NaH}_2\text{PO}_4 \cdot 2\text{H}_2\text{O}) = \frac{0.31}{156}$
$n(\text{Na}_2\text{HPO}_4 \cdot 12\text{H}_2\text{O}) = 7.374 \times 10^{-3}$	$n(\text{NaH}_2\text{PO}_4 \cdot 2\text{H}_2\text{O}) = 1.987 \times 10^{-3}$

$$C = \frac{n}{V}$$

$C(\text{Na}_2\text{HPO}_4 \cdot 12\text{H}_2\text{O}) = \frac{7.374 \times 10^{-3}}{1}$	$C(\text{NaH}_2\text{PO}_4 \cdot 2\text{H}_2\text{O}) = \frac{1.987 \times 10^{-3}}{1}$
$C = 7.37 \times 10^{-3} \text{ mol L}^{-1}$	$C = 1.99 \times 10^{-3} \text{ mol L}^{-1}$

Using K_a

$$K_a = \frac{[\text{HPO}_4^{2-}][\text{H}_3\text{O}^+]}{[\text{H}_2\text{PO}_4^-]}$$

$$6.3 \times 10^{-8} = \frac{7.37 \times 10^{-3} [\text{H}_3\text{O}^+]}{1.99 \times 10^{-3}}$$

$$1.703 \times 10^{-8} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log 1.703 \times 10^{-8}$$

$$\text{pH} = 7.768$$

Answer: 7.77

OR

alternative method using H-H equation

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log (6.3 \times 10^{-8})$$

$$\text{p}K_a = 7.2$$

$$\text{pH} = 7.2 + \log \frac{7.37 \times 10^{-3}}{1.99 \times 10^{-3}}$$

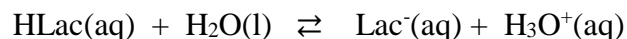
$$\text{pH} = 7.2 + \log 3.703$$

$$\text{pH} = 7.2 + 0.5686$$

$$\text{pH} = 7.768$$

Answer: 7.77

50mL of 0.15 mol L⁻¹ Lactic acid (HLac) is mixed with 35mL of 0.25 mol L⁻¹ Sodium lactate (Lac⁻) what is the pH of the buffer solution?
 $K_a(\text{HLac}) = 1.38 \times 10^{-4}$



final concentration = $\frac{\text{original concentration} \times \text{original volume}}{\text{final volume}}$

$C(\text{HLac}) = \frac{0.15 \times 0.05}{0.085}$	$C(\text{Lac}^-) = \frac{0.25 \times 0.035}{0.085}$
$C(\text{HLac}) = 0.0882 \text{ mol L}^{-1}$	$C(\text{Lac}^-) = 0.1029 \text{ mol L}^{-1}$

Using K_a

$$K_a = \frac{[\text{Lac}^-][\text{H}_3\text{O}^+]}{[\text{HLac}]}$$

$$1.38 \times 10^{-4} = \frac{0.1029 [\text{H}_3\text{O}^+]}{0.0882}$$

$$1.182 \times 10^{-4} = [\text{H}_3\text{O}^+]$$

$$\text{pH} = -\log 1.182 \times 10^{-4}$$

$$\text{pH} = 3.926$$

Answer: 3.93

OR

alternative method using H-H equation

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log (1.38 \times 10^{-4})$$

$$\text{p}K_a = 3.86$$

$$\text{pH} = 3.86 + \log \frac{0.1029}{0.0882}$$

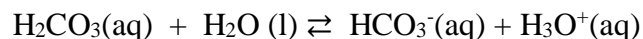
$$\text{pH} = 3.86 + \log 1.166$$

$$\text{pH} = 3.86 + (0.0669)$$

$$\text{pH} = 3.926$$

Answer: 3.93

Calculate the ratio of carbonic acid concentration to bicarbonate ion concentration buffer in the blood which maintains a pH of 7.4.
 $K_a(\text{H}_2\text{CO}_3) = 4.3 \times 10^{-7}$



$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{inverse log } -7.4 = [\text{H}_3\text{O}^+]$$

$$3.98 \times 10^{-8} = [\text{H}_3\text{O}^+]$$

Using K_a

$$K_a = \frac{[\text{HCO}_3^-][\text{H}_3\text{O}^+]}{[\text{H}_2\text{CO}_3]}$$

$$4.3 \times 10^{-7} = \frac{[\text{HCO}_3^-][3.98 \times 10^{-8}]}{[\text{H}_2\text{CO}_3]}$$

$$\frac{4.3 \times 10^{-7}}{3.98 \times 10^{-8}} = \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$10.8 = \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$[\text{H}_2\text{CO}_3]$$

$$\text{Answer: } [\text{HCO}_3^-] : [\text{H}_2\text{CO}_3] = 11 : 1$$

OR

alternative method using H-H equation

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log 4.3 \times 10^{-7}$$

$$\text{p}K_a = 6.366$$

$$7.4 = 6.366 + \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$7.4 - 6.366 = \log \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

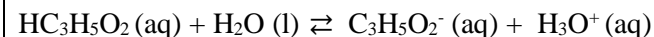
$$\text{inverse log } 1.033 = \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$10.78 = \frac{[\text{HCO}_3^-]}{[\text{H}_2\text{CO}_3]}$$

$$[\text{HCO}_3^-] : [\text{H}_2\text{CO}_3] = 11 : 1$$

$$\text{Answer: } [\text{HCO}_3^-] : [\text{H}_2\text{CO}_3] = 11 : 1$$

Calculate the concentration of $\text{C}_3\text{H}_5\text{O}_2^-$ in a buffer solution which has a pH of 4.5 if the concentration of $\text{HC}_3\text{H}_5\text{O}_2$ is 0.50 mol L⁻¹
 $K_a(\text{HC}_3\text{H}_5\text{O}_2) = 1.34 \times 10^{-5}$



$$\text{pH} = -\log [\text{H}_3\text{O}^+]$$

$$\text{inverse log } -4.5 = [\text{H}_3\text{O}^+]$$

$$3.16 \times 10^{-5} = [\text{H}_3\text{O}^+]$$

Using K_a

$$K_a = \frac{[\text{C}_3\text{H}_5\text{O}_2^-][\text{H}_3\text{O}^+]}{[\text{HC}_3\text{H}_5\text{O}_2]}$$

$$1.34 \times 10^{-5} = \frac{[\text{C}_3\text{H}_5\text{O}_2^-] 3.16 \times 10^{-5}}{0.50}$$

$$\frac{(1.34 \times 10^{-5})(0.50)}{3.16 \times 10^{-5}} = [\text{C}_3\text{H}_5\text{O}_2^-]$$

$$0.2120 = [\text{C}_3\text{H}_5\text{O}_2^-]$$

Answer 0.212 mol L⁻¹

OR

alternative method using H-H equation

$$\text{p}K_a = -\log K_a$$

$$\text{p}K_a = -\log (1.34 \times 10^{-5})$$

$$\text{p}K_a = 4.872$$

$$4.5 = 4.872 + \log \frac{[\text{C}_3\text{H}_5\text{O}_2^-]}{0.50}$$

$$4.5 - 4.872 = \log \frac{[\text{C}_3\text{H}_5\text{O}_2^-]}{0.50}$$

$$-0.372 = \log \frac{[\text{C}_3\text{H}_5\text{O}_2^-]}{0.5}$$

$$\text{inverse log } (-0.372) = \frac{[\text{C}_3\text{H}_5\text{O}_2^-]}{0.5}$$

$$(0.4246)(0.5) = [\text{C}_3\text{H}_5\text{O}_2^-]$$

$$0.2120 = [\text{C}_3\text{H}_5\text{O}_2^-]$$

Answer 0.212 mol L⁻¹